9) **Electrochemistry -** a branch of chemistry dealing with the relations between chemical energy and electrical energy.

a) Examples of technology utilizing electrochemistry:
   i) Batteries
   ii) Electroplating

11) **Oxidation -** chemical process where electrons are lost. The oxidation number (charge) gets more positive.

12) **Reduction -** chemical process where electrons are gained. The oxidation state (charge) gets more negative.

13) **OIL RIG -** an acronym to help remember the difference between oxidation and reduction. Oxidation Is Loss, Reduction Is Gain (of electrons).

14) **Electrochemical cell -** a system consisting of electrodes that dip into an electrolytic solution and in which a chemical reaction produces an electric current.

a) **Electrolytic solution -** an aqueous solution that contains ions (electrolytes) and conducts electric current.
   i) Examples:
      (1) "Gatorade", "All-sport" …
      (2) "Battery acid" (although it is a base in alkaline batteries).
      (3) Your blood and interstitial fluid.

b) There are two types of electrochemical cells:
   i) **Voltaic cell -** uses energy from a spontaneous chemical reaction to produce an electric current. These cells are often called a "Galvanic cells".
      (1) Examples:
         (a) A non-rechargeable battery ("Dry cell")
         (b) A car battery
         (c) A potato clock
         (d) A "plating solution and cell" that does not require additional electricity. (We will copper plate tin this way in the lab.)
      (2) **Anode -** is always the electrode where oxidation occurs. It is the negative terminal in a voltaic cell.
      (3) **Cathode -** is always the electrode where reduction occurs. It is the positive terminal in a voltaic cell.
      (4) Diagrams on following pages.

ii) **<u>Electrolytic cell -</u>** uses energy from a non-spontaneous chemical reaction to produce an electric current. These are often used to plate metals.

(1) Examples:

(a) Electrolysis of water into Hydrogen and Oxygen.

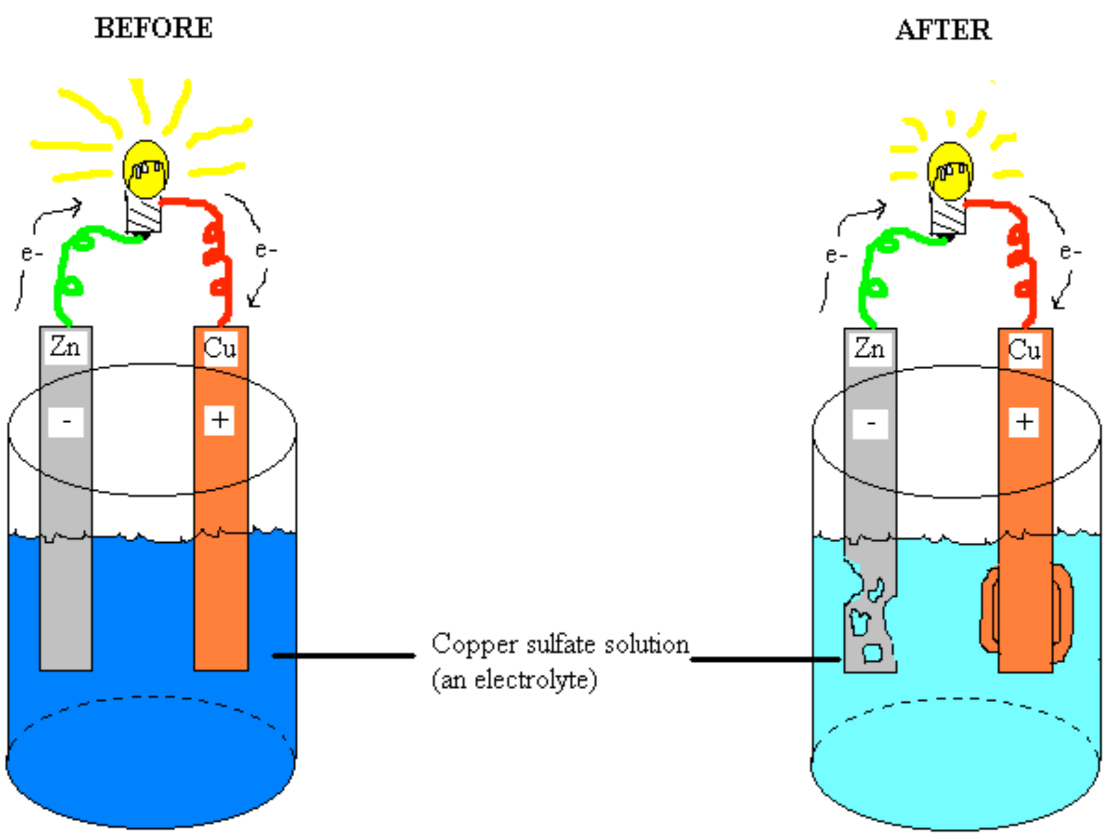
(b) Electrolysis of brine (concentrated salt water) into Chlorine gas, Hydrogen, and Sodium chloride.

(2) **<u>Anode -</u>** <u>is always</u> the electrode where oxidation occurs. It is the positive terminal in an electrolytic cell.

(3) **<u>Cathode -</u>** <u>is always</u> the electrode where reduction occurs. It is the negative terminal in an electrolytic cell.

(4) Diagrams on following pages.

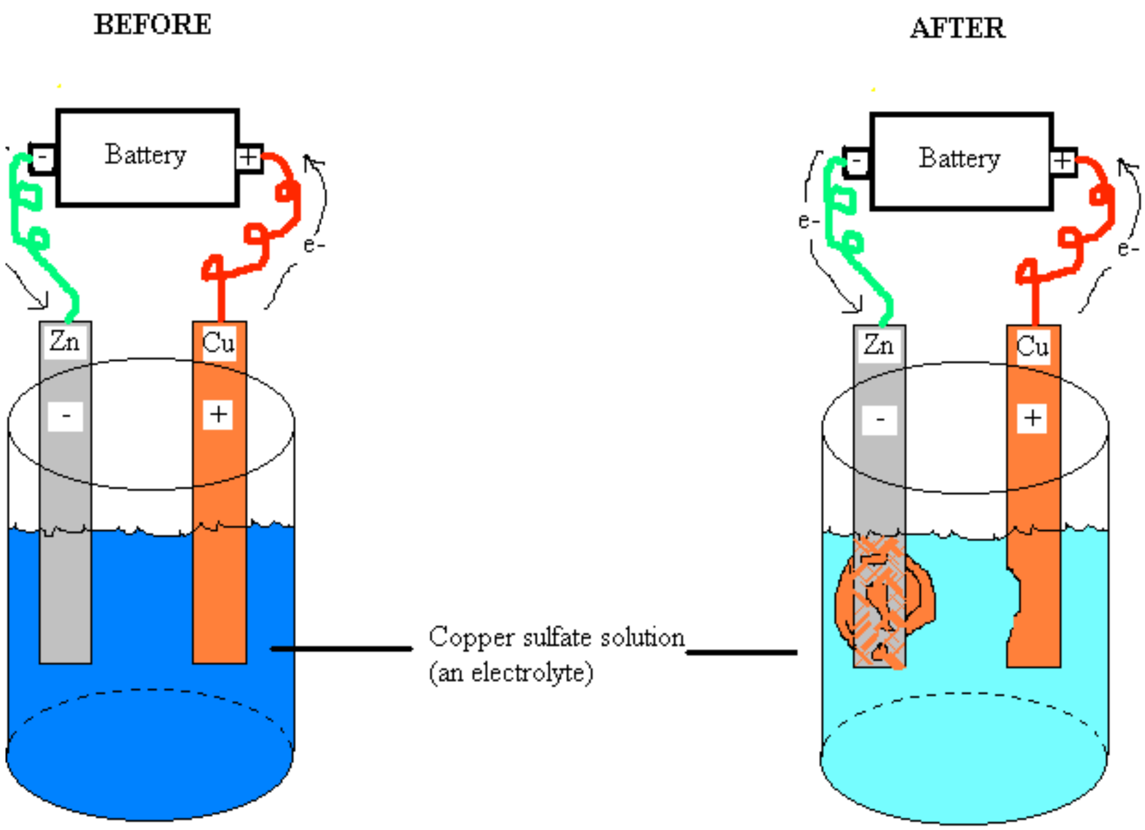
Diagram of a voltaic cell.



Notice that as time goes on the copper will spontaneously plate on to the copper cathode. Simultaneously, the zinc (which is loosing

electrons and putting Zinc cations into solution) will become severely corroded, causing the voltage to decrease and the light to dim. Copper when in solution creates a blue liquid. As the copper leaves the solution by plating onto the copper strip the color fades. This is essentially what happens when a battery "dies".

Diagram of an electrolytic cell.



Notice how the external power source forces the electrons to travel opposite their spontaneous direction. This time the copper is oxidized and thus corroded (albeit after the majority of copper cations in solution have plated the tin). The tin is plated with the copper. Once the external power source is removed and the "light" is replaced, this cell will once again become a voltaic cell.

15) **Half-cell reaction -** the chemical reaction that occurs at one of the electrodes in an electrolytic cell. It involves the electrode and electrons only.

a) If oxidation is occurring, the electrode is losing electrons.

i) $Cu \rightarrow\rightarrow\rightarrow Cu^{+2} + 2e^{-}$

(1) In this example copper metal will be converted into copper ions. (Often a blue liquid)

b) If reduction is occurring, the electrode is gaining electrons.

i) Zn+2 + 2e- →→→ Zn

(1) In this example zinc ions in solution will be converted into zinc metal.

16) **Half-cell potential -** the sum of the voltages (both positive and / or negative from the oxidation and reduction reactions that occur in an electrolytic cell.

a) **Standard reduction potential -** a measure of the voltage gained or lost when an electrode is reduced in an electrochemical cell.

b) **Cell potential -** the sum of the standard reduction potential and the standard oxidation potential (S.R.P. with the opposite charge.)

i) Cell potentials are always calculated from the standard reduction potential.

ii) The electrode that is being oxidized will have the same value, but the opposite charge; of what is listed as the S.R.P..

iii) Example: 1 Using the metals pictured in the previous diagrams: Cu & Zn.

Calculate the cell potential based on their standard reduction potentials.

(1) STEP 1: Write down the S.R.P. of the metals involved. (You can find these listed in a table in any chemistry text book.) For these elements they are:

$Cu^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Cu \quad = +0.34$ volt
$Zn^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Zn \quad = -0.76$ volt

(2) STEP 2:

(a) Rewrite the more positive voltage and the corresponding chemical equation. (Since it is the most positive, we will not change it; and it will be the electrode where *Reduction* occurs.)

$Cu^{+2} + 2e^{-} \rightarrow\rightarrow\rightarrow Cu \quad = +0.34$ volt

(b) Write the opposite of the remaining voltage and its corresponding chemical equation. (For the voltage, all you have to do is switch the "sign" of the charge. For the equation: you will switch everything from the left side of the arrow to the right side, and everything from the right side to the left side.)

$Zn \rightarrow\rightarrow\rightarrow Zn^{+2} + 2e^{-} \quad = +0.76$ volt

(3) STEP 3:
Determine if the electrons on both sides of the equation cancel each other out. If they do not, you will have to multiply one or both of the half-cell reactions by a whole number until they do cancel. (What ever number you choose to multiply the reaction equation by must also be multiplied with the voltage.) As you will see they do cancel in this example.

$Cu^{+2} + \mathbf{2e^{-}} \rightarrow\rightarrow\rightarrow Cu \quad = +0.34$ volt
$Zn \rightarrow\rightarrow\rightarrow Zn^{+2} + \mathbf{2e^{-}} \quad = +0.76$ volt

(4) STEP 3:
Calculate the cell potential by adding the voltages determined above together. (This is how much voltage you should get from the two metals in a 1<u>M</u> solution of electrolyte)

$(+0.34$ volt$)$
$\underline{+(+0.76 \text{ volt})}$
+1.10 volt = cell potential **(Since this is a positive value this would be a spontaneous reaction, and thus a voltaic cell.)**

iv) Example 2: Using the metals Zn^{+2}/Zn & Cr^{+3}/Cr Calculate the cell potential based on their standard reduction potentials.

(1) STEP 1: Write down the S.R.P. of the metals involved. (You can find these listed in a table in any chemistry text book.) For these elements they are:

$Cr^{+3} + 3e^- \rightarrow\rightarrow\rightarrow Cr$ = -0.74 volt
$Zn^{+2} + 2e^- \rightarrow\rightarrow\rightarrow Zn$ = -0.76 volt

(2) STEP 2:

(a) Rewrite the more positive voltage and the corresponding chemical equation. (Since it is the most positive, we will not change it; and it will be the electrode where *Reduction* occurs.)

$Cr^{+3} + 3e^- \rightarrow\rightarrow\rightarrow Cr$ = -0.74 volt

(b) Write the opposite of the remaining voltage and its corresponding chemical equation. (For the voltage, all you have to do is switch the "sign" of the charge. For the equation: you will switch everything from the left side of the arrow to the right side, and everything from the right side to the left side.)

$Zn \rightarrow\rightarrow\rightarrow Zn^{+2} + 2e^-$ = +0.76 volt

(3) STEP 3:

Determine if the electrons on both sides of the equation cancel each other out. If they do not, you will have to multiply one or both of the half-cell reactions by a whole number until they do cancel. (What ever number you choose to multiply the reaction equation by must also be multiplied with the voltage.) As you should notice, they do not cancel in this example.

Cr^{+3} + **$3e^-$** →→→ Cr = -0.74 volt
Zn →→→ Zn^{+2} + **$2e^-$** = +0.76 volt

The smallest common denominator is 6 in this example. We will have to multiply the Chromium by 2; and the Zinc by 3.

2Cr^{+3} + **$6e^-$** →→→ **2**Cr = **-1.48** volt
3Zn →→→ **3**Zn^{+2} + **$6e^-$** = **+2.28** volt

(4) STEP 3:
Calculate the cell potential by adding the voltages determined above together. (This is how much voltage you should get from the two metals in a 1M solution of electrolyte)

(-1.48 volt)
+(+2.28 volt)
+0.80 volt = cell potential **(Again, since this is a positive value this would be a spontaneous reaction, and thus a voltaic cell. If the calculated value would be negative, this cell would not be spontaneous and the reaction would require an external power source. Thus, if negative, it would be electrolytic.)**